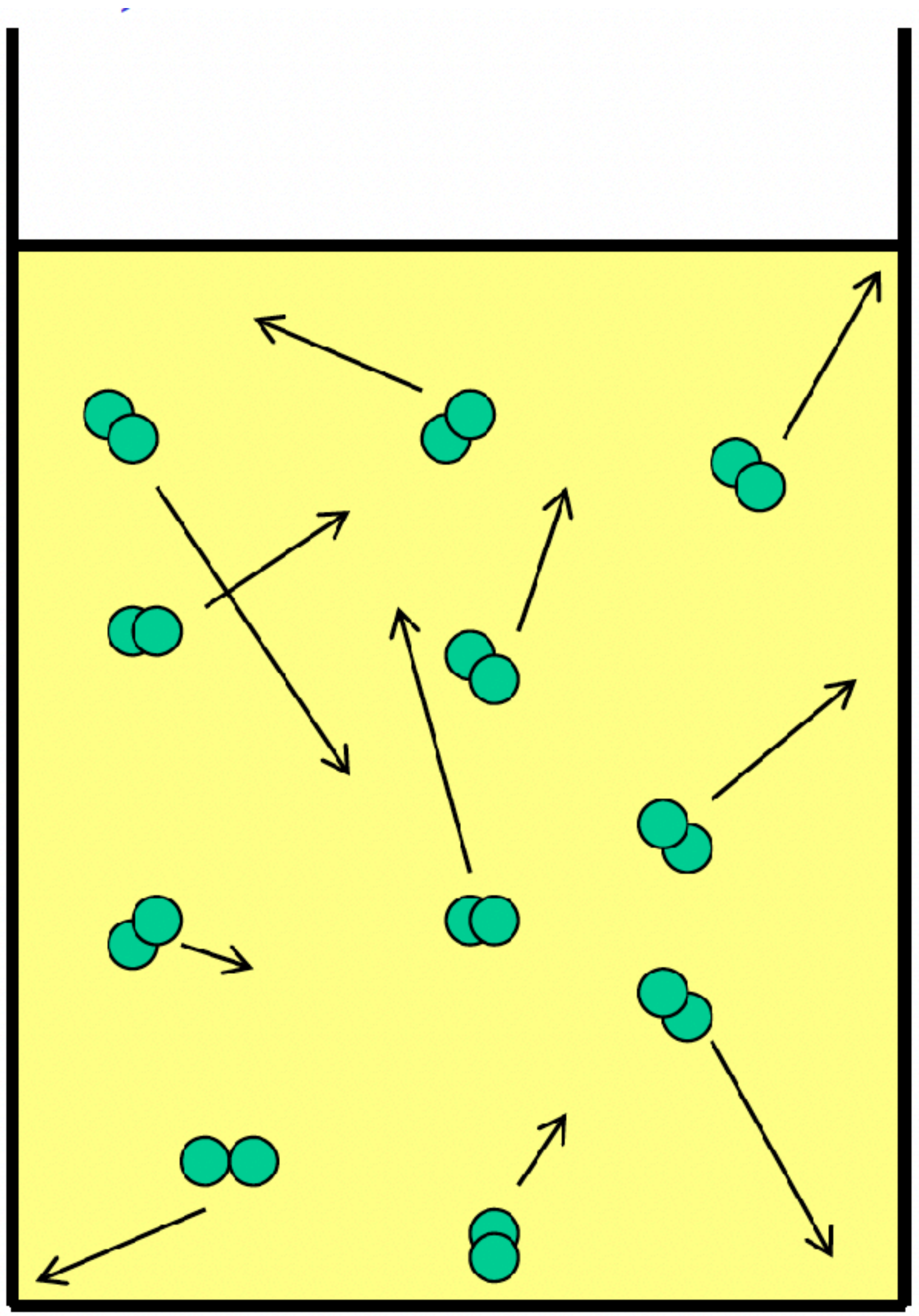
3.5 Kinetic Molecular Theory

- KMT
- Maxwell-Boltzmann Distribution
- Kinetic Energy

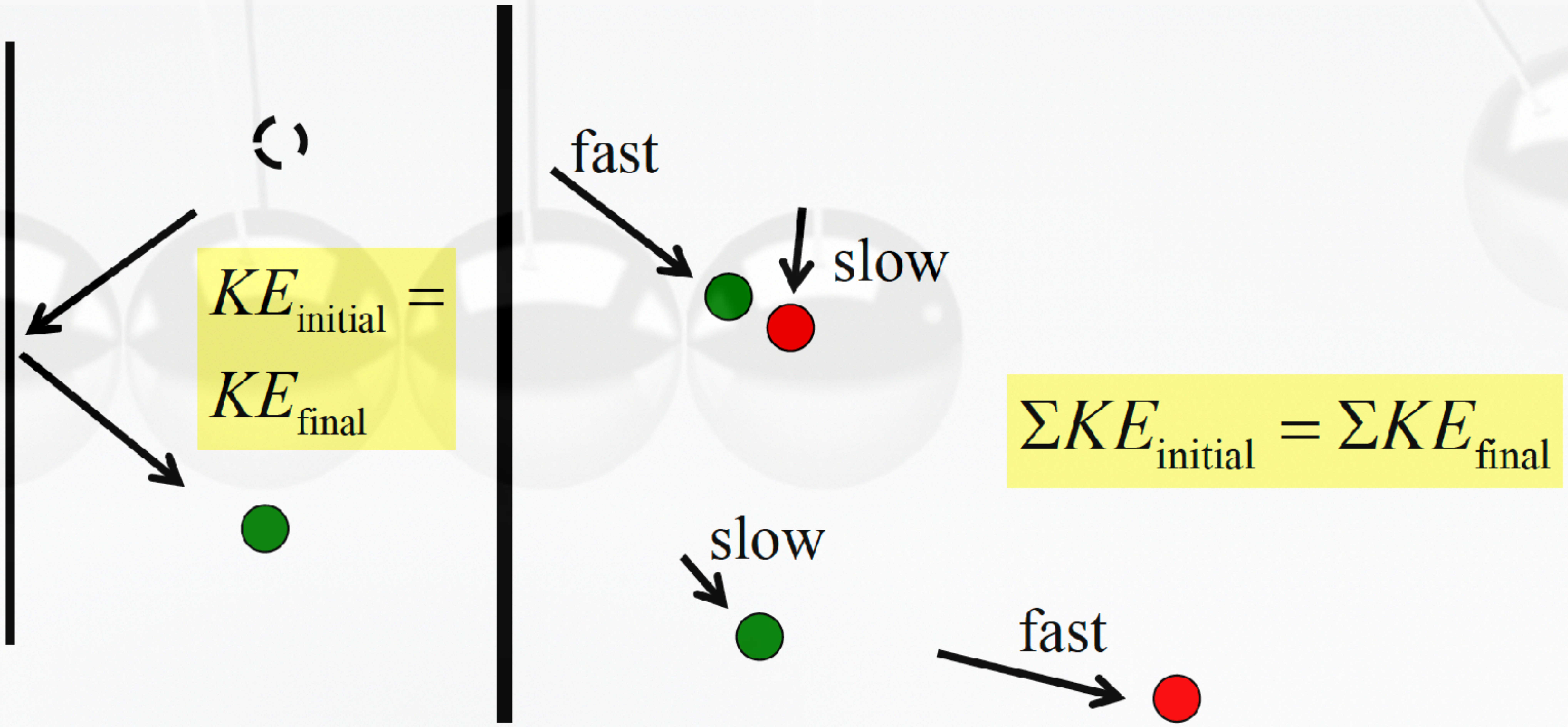
Kinetic Molecular Theory

- Gases consist of particles (molecules and/or individual atoms) that are in continuous random movement.
- Total volume is negligible when compared to the volume of the system.
- Ideal Gas Law assumes that the volume of the gas particles in a system is zero.
- Coulombic forces of attraction or repulsion do not exist between gas particles.



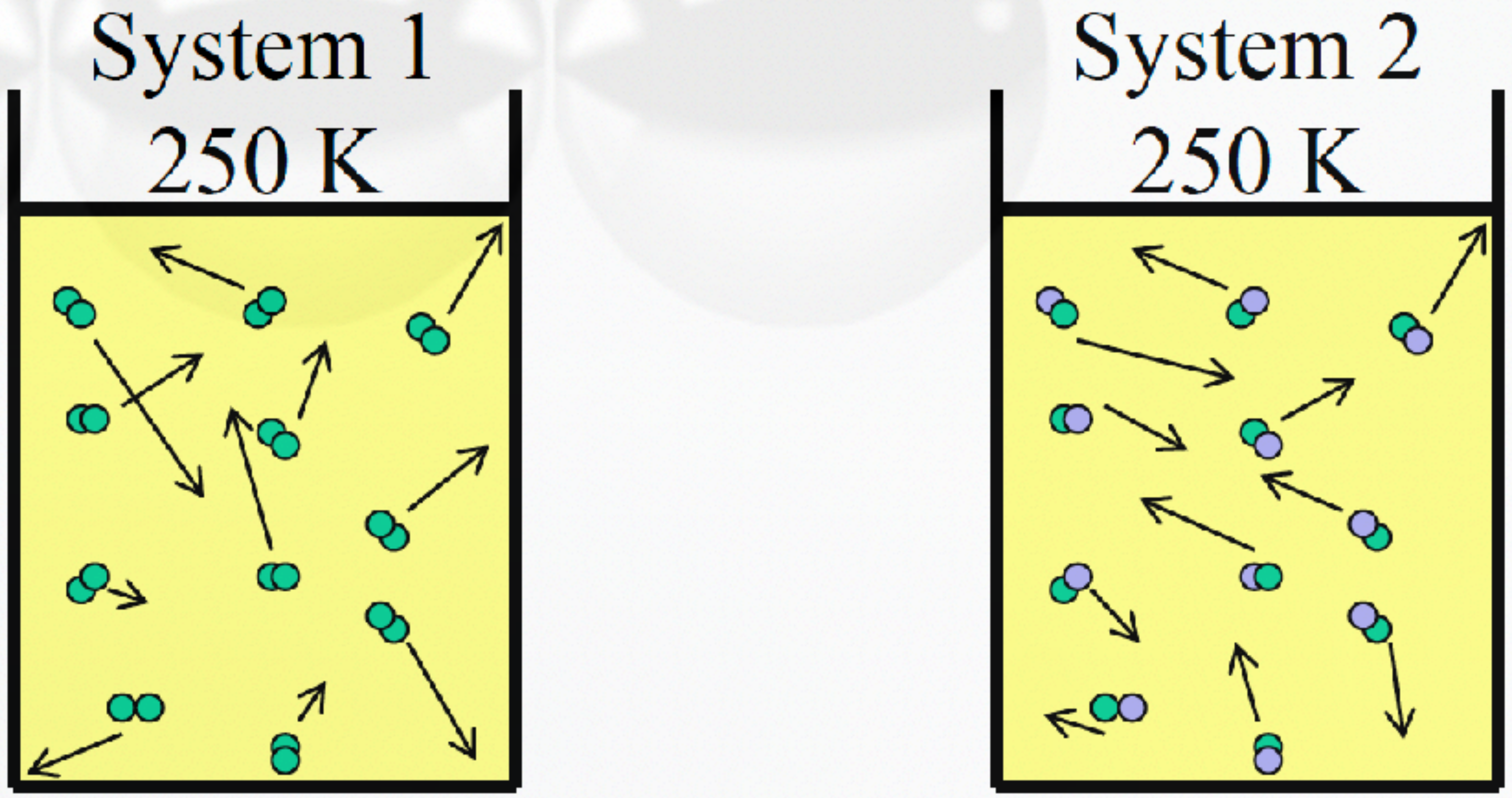
Kinetic Molecular Theory

- Collisions experienced by gas particles are elastic (*Kinetic Energy is Conserved.)*



Kinetic Molecular Theory

- The average KE of the gas particles in a system is proportional to the absolute temperature.
- The gas particles in any system that is kept at the same temperature will have the same average KE.
    - Average KE in system 1 = average KE in system 2.

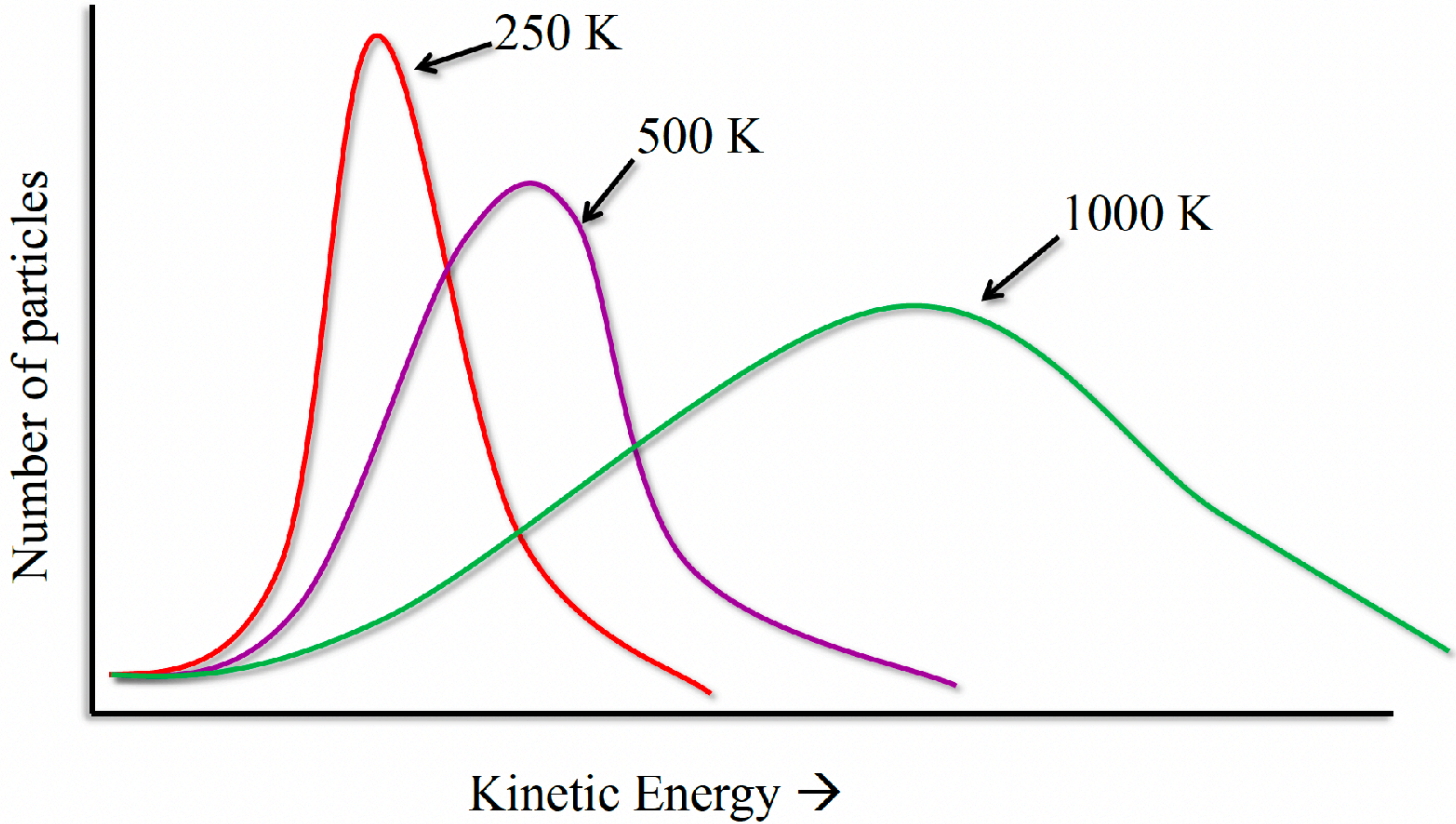


Kinetic Energy of Gas Molecules

- Translational Energy
  - Gas molecules move through space in straight lines (no attractive forces)
  - Rotational Energy
  - Vibrational Energy

Most of a gas particle's KE is related to its translational velocity.

Maxwell-Boltzmann Distribution, Temp & Pressure



Maxwell-Boltzmann Distribution, Temp & Pressure

- The average KE of the particles in a system increases as the temperature increases.
  - At any temperature, there is a large range of kinetics.
- At any given temperature, the particles with less KE exert a lower pressure and the particles with more KE exert a higher pressure.
  - The total pressure exerted by the gas particles in a system is an average.

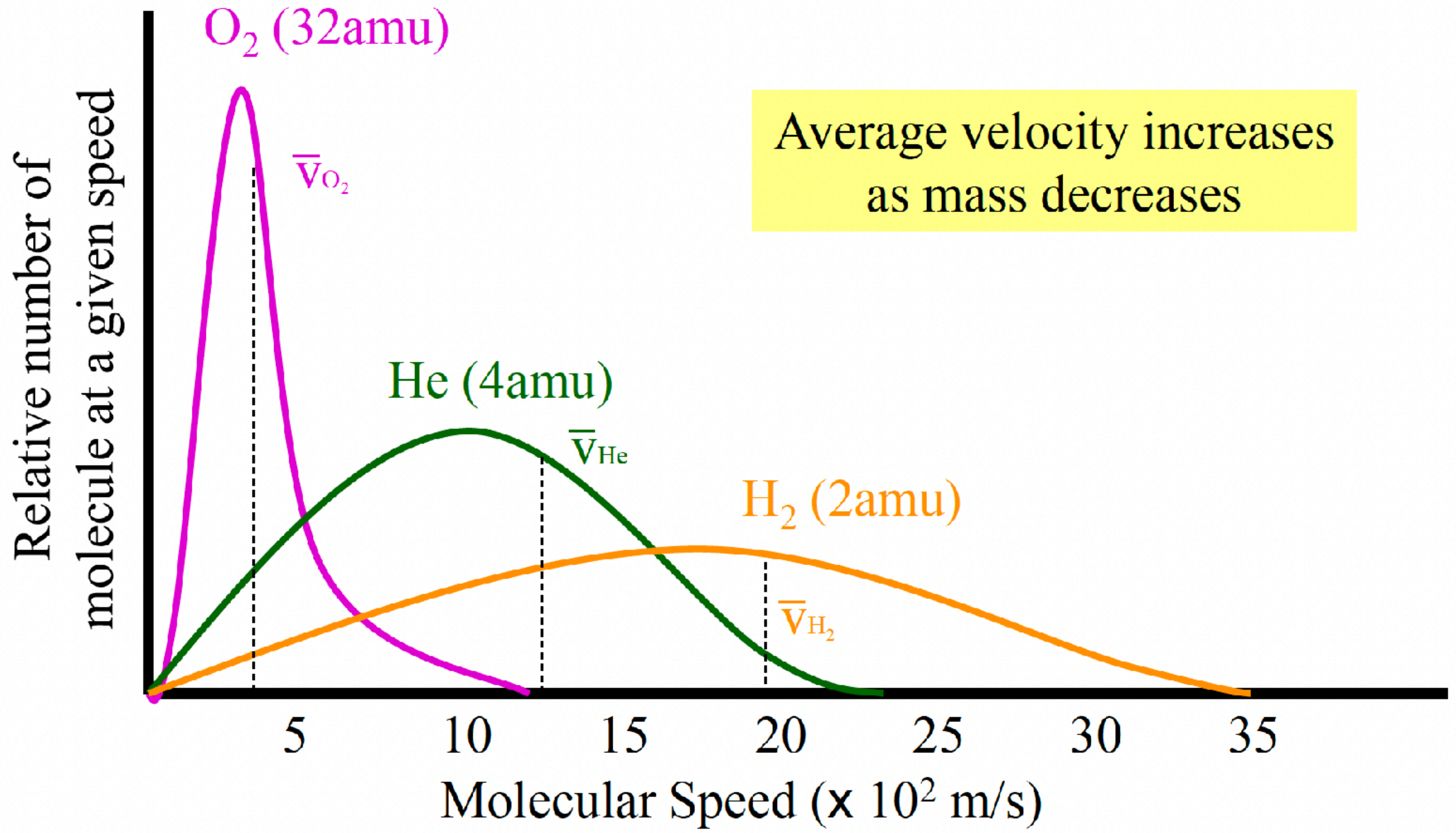
KE, Mass & Velocity of a Single Gas Particle

$$KE = 1/2\ mv^2$$

v = velocity of a specific gas particle (m/s)

m = mass of that particle (kg)

Molar Mass & Molecular Speed (@25˚C)



Molar Mass & Molecular Speed & Temperature

- At any given temperature, the average *KE* of all gas particles is the same.
- Gases with smaller molar masses will have higher average velocities.
- Gases with larger molar masses will have lower average velocities.

$$KE = 1/2\ mv^2$$

3.6 Deviation from Ideal Gas Law

- Real vs. Ideal Gases

All Real Gases <u>Do Not</u> Behave Ideally When…

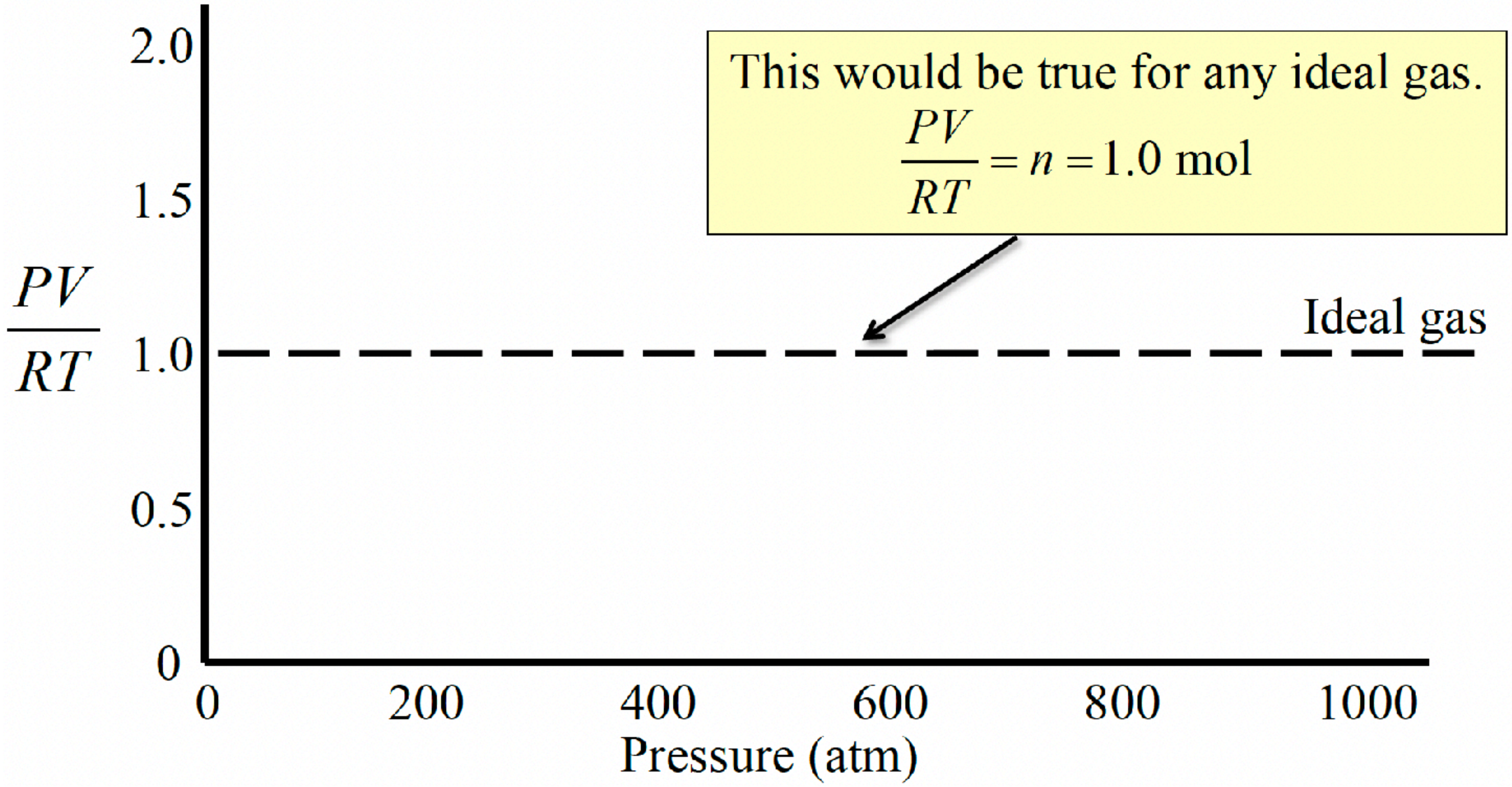
- Under high pressures (P > 5 atm)
- At low temperatures

Under such conditions, the ideal gas equation

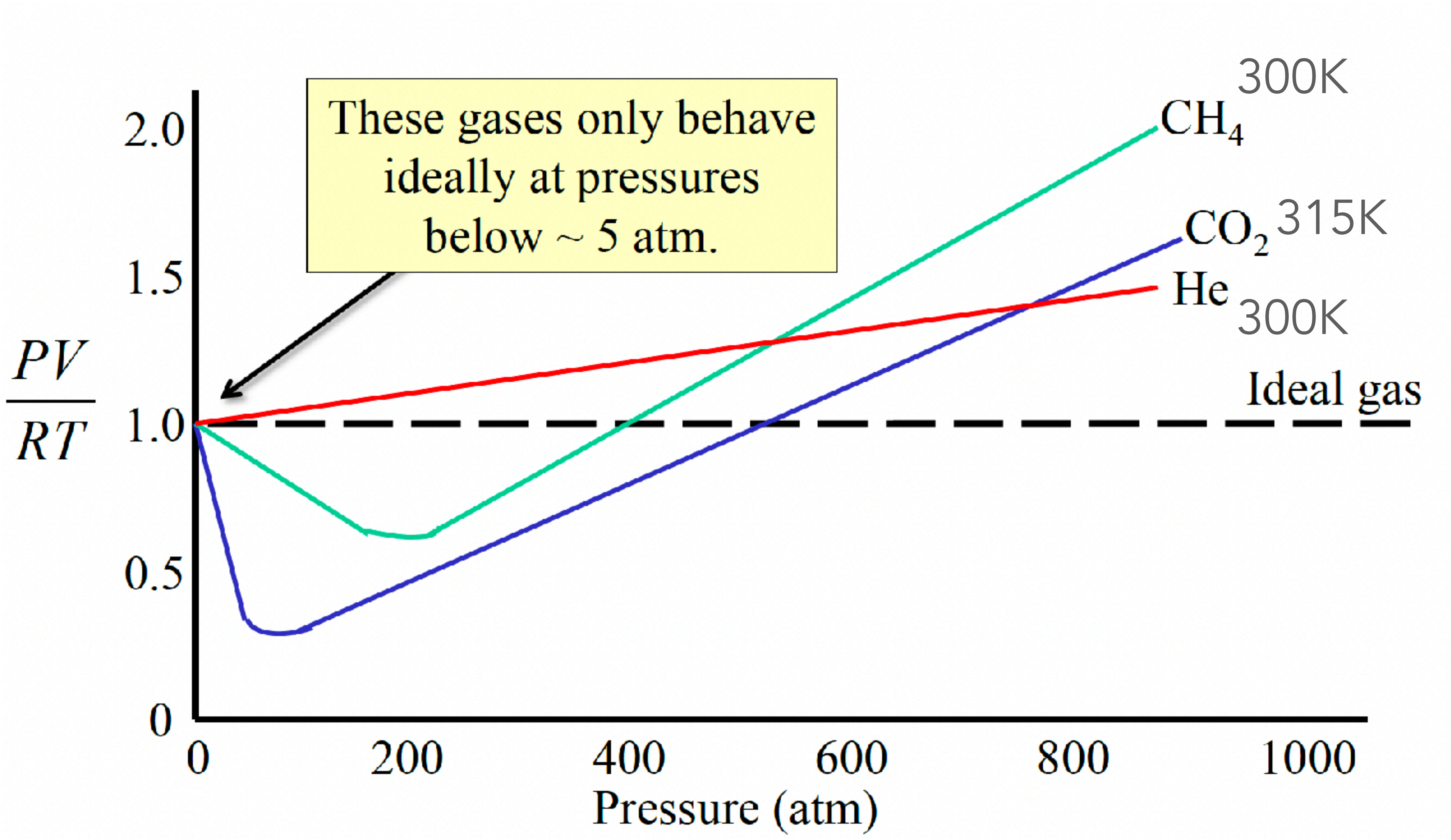
$$PV = nRT$$

<u>does not</u> make accurate predictions.

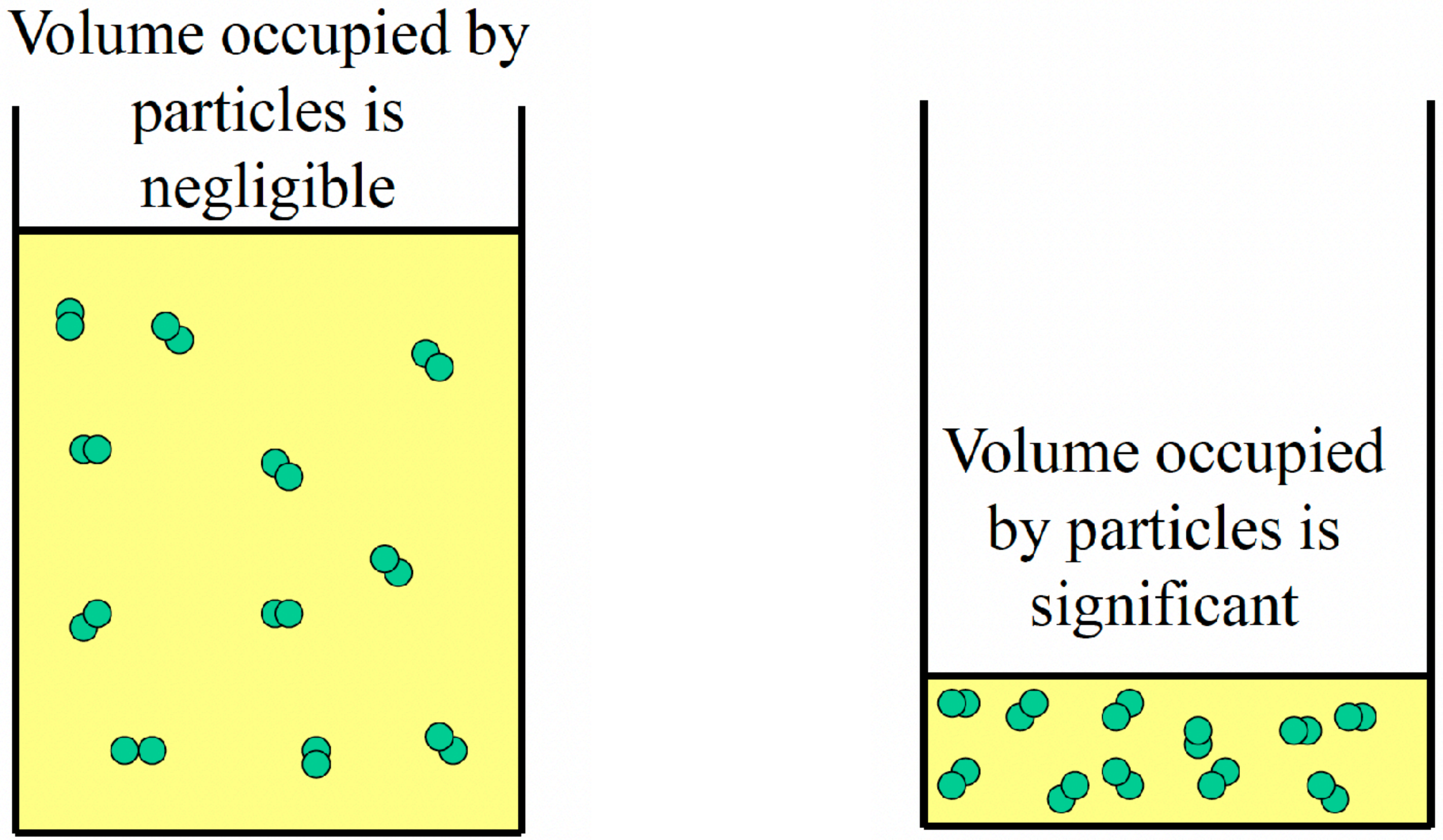
PV/RT vs. P for 1.0 mole of Ideal Gas



PV/RT vs. P for 1.0 mole of Different Gases at Constant T



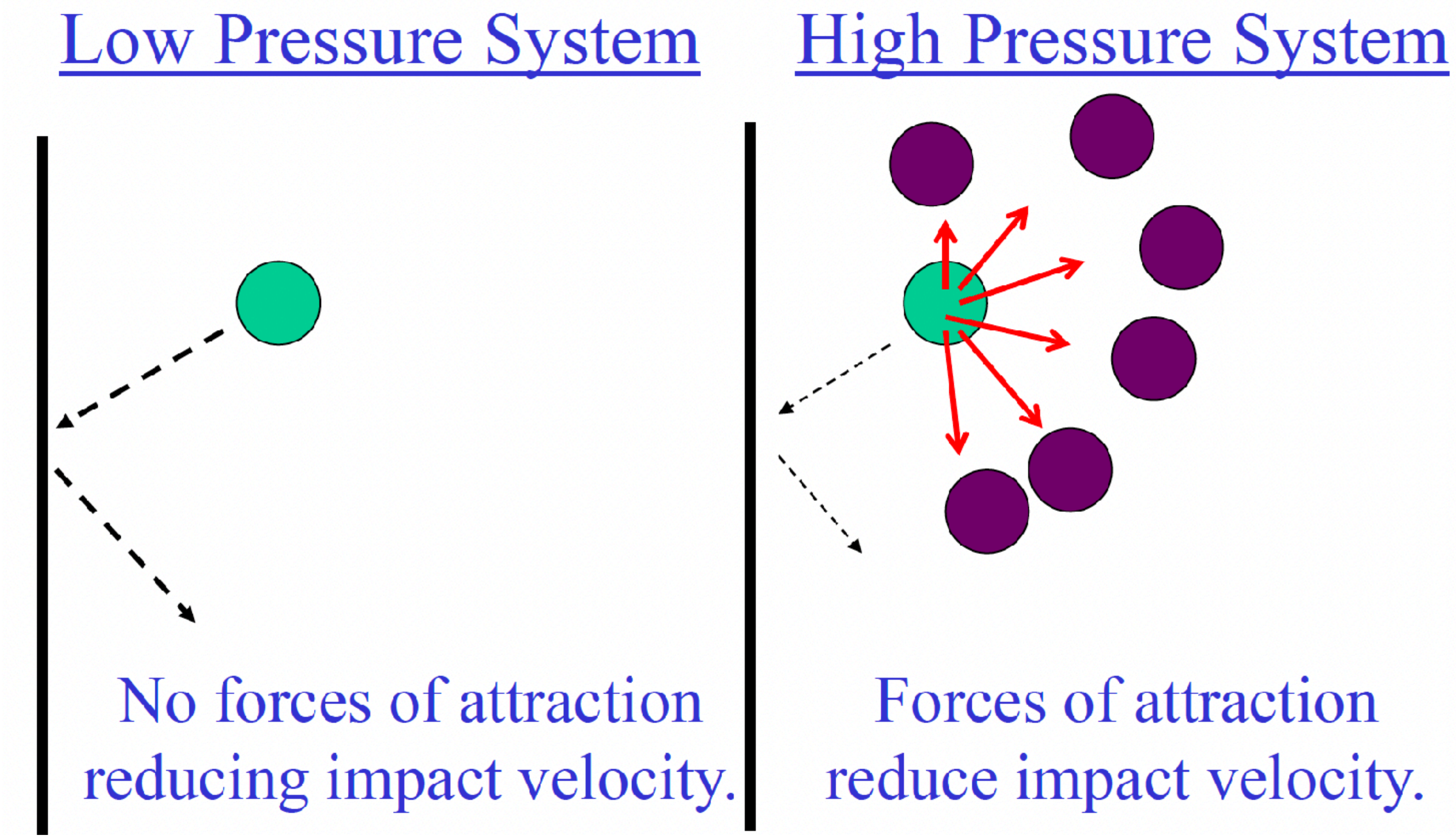
Volume Adjustment for Gases Under High Pressure



Pressure Adjustment for Gases Under High Pressures (low volume)

- When gas particles are very close together, they pressure they exert may be less than what the Ideal Gas equation would predict.
- Neighboring molecules exert forces of attraction on one another when they are very close together.
- Such forces pull a gas molecule in the direction opposite to its motion.
- This reduces the pressure resulting from impacts with the walls of the container.

Pressure Adjustment for Gases Under High Pressures (low volume)



Van der Waals Equation

$$\left(P + \frac{n^2 a}{V^2}\right)(V - nb) = nRT$$

P = actual or measured pressure (atm)

n = moles of gas

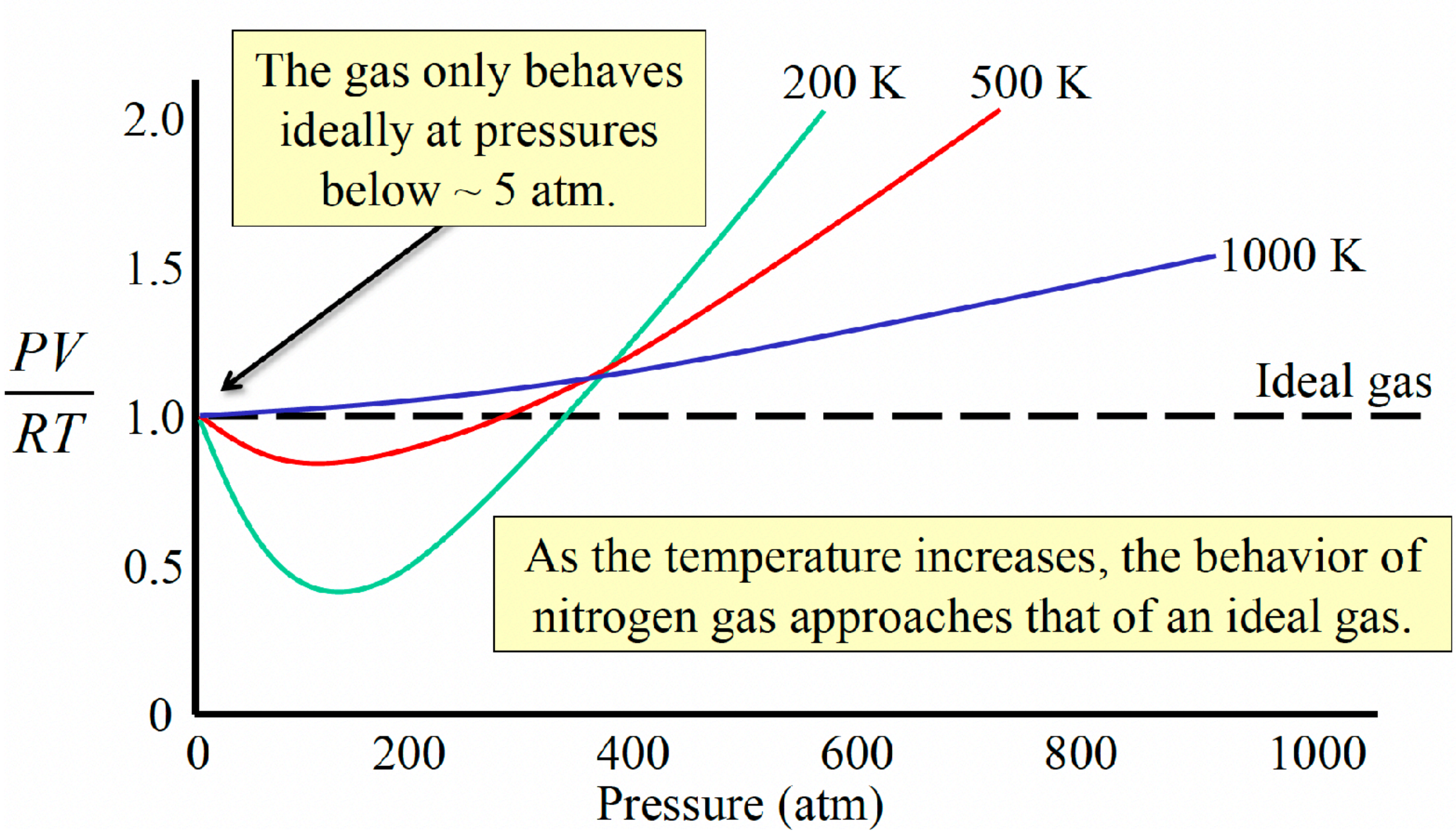
a and b = constants for the specific gas in question

V = actual or measured volume (L)

T = temperature (K)

R = 0.0821 L•atm/mol•K

PV/RT vs. P for 1.0 mole of $N_2(g)$ at Different Temperatures



Gases do not Behave Ideally at Low Temperatures

- The Ideal Gas law assumes that gases experience no intermolecular forces of attraction.
- At high temperatures, the kinetic energy of gas particles overcome any intermolecular forces of attraction.
- At low temperatures, gas particles move slower and are closer together. Attractions between molecules exist under these conditions.

Non-Ideal Behavior & Condensation

- IMFs increase as the distance between particles decreases.
- Can lead to condensation at low T and high P.
- This applies to all gases.